REMARKS

This amendment is responsive to the Final Office Action dated August 4, 2005. Claims 1, 7, 9, 13, 15, 17, 21, 23, 25, 27, 29, 33, 39, 41, 45, and 54 are pending in this application and have been rejected. These claims have been indicated to be allowable if rewritten or amended to overcome rejections under 35 USC § 112 (second paragraph). Reexamination is respectfully requested.

These remarks follow the order of the outstanding Office Action beginning at page 2 thereof.

Drawings

The Examiner has objected to the drawings because of the reference characters (61) and (62) which the Examiner asserts designate both incoming terminals and anode terminals. In response, Applicant has amended the specification at page 44 to agree with the drawings. As stated at line 11, the incoming terminals are terminals (61), (62), (63) as shown in the drawing. However, the incoming terminals comprise both cathode terminals and anode terminals. In order to make this clear, Applicant has specified terminal (63) as being an incoming terminal and terminal (61) and (62) as being incoming terminals. It is noted that (63) is a cathode terminal and terminals (61) and (62) are anode terminals. In the electrical art, it is well known that terminals may be of either the anode type or the cathode type, and that they may be incoming terminals. The drawing, on the other hand, shows clearly that there are two different kinds of

terminals, namely terminal (63) (cathode) which is different from terminals (62) and (61) which are the anode terminals. It is respectfully submitted that this amendment to the specification does not add new matter since the cathode terminal as originally stated in line 19 is designated as serving as an incoming terminal and that the anode terminals (61) and (62) are designated in line 21 as serving as incoming terminals. This is wholly consistent with line 11 which states that all of terminals (61), (62) and (63) are incoming.

Specification

The Examiner has objected to the amendment filed May 26, 2005 on the grounds that it introduces new matter. Applicant respectfully submits that this amendment should be entered.

Attached hereto as Exhibit 1 is a derivation of formula (3). Formula (3) is the conversion between energy in electron volts and wavelength. This in turn is controlled by three constants which are the Plank's Constant, the speed of light, and the energy of an electron volt (see Equation (2)).

In Exhibit 1, page 2 there is set forth a table which is the same as the first two columns of Tables 3 and 5 which appear in the specification at pages 35 and 39. The difference between the tables appearing in the specification and the table in Exhibit 1 is that the tables in the specification have the energy value in electron volts rounded back to two places. If the physics formula is followed precisely, the energy values will appear as Applicant shows them in the right hand column of Exhibit 1 where the values are

extended out to six places. This attached table shows that when Formula (3) using the constant of 1240 is used, that the proper value for the electron voltage in each table is 2.19 for the first value and 2.16 for the second value.

Although Applicant has demonstrated that the equation for conversion between wavelength and energy is derived from Plank's Constant, Applicant also submits as Exhibit 2 an explanation of the fundamental physics of light which demonstrates in the equation at page 3 that the energy of the photon in electron volts (eV) is equal to 1240 eV – nm/wavelength in nm.

Applicant, therefore, submits that the error in Tables 3 and 5, items 1 and 2 is a clear mathematical error and that the fundamental physics which dictate that the conversion between column 1 and column 2 of these tables is clearly 1240. Still further, Applicant has demonstrated in Exhibit 1 that this is true for every case in Tables 3 and 5, entries 3 – 11. Still further, the Examiner should note that in Table 6, the same rule applies, namely the conversion from energy in electron volts to wavelength is once again 1240.

Therefore, the correction is merely correction of a mathematical mistake, not new matter. Should the Examiner assert that it is new matter, then the Patent Office will be taking a position contrary to fundamental physics and Plank's Constant.

Applicant submits that this new evidence clearly supports the amendment.

<u>Claim Rejections – 35 USC § 112</u>

Applicant has amended claim 1 to provide a clearer definition of the invention.

This definition of the invention as now stated in claim 1 is supported by Applicant's specification, page 29, lines 20 – 23 and page 30, lines 15 – 18. It is respectfully submitted that this proposed modification in claim 1 overcomes the rejection found at paragraph 6, page 3 of the outstanding Office Action. However, should the Examiner still find this claim to be objectionable, it is respectfully requested that the Examiner telephone the undersigned at 202-347-2600 in order to resolve this matter and provide appropriate language which is acceptable.

In view of the foregoing, it is respectfully submitted that the application is now in condition for allowance, and early action in accordance thereof is requested. In the event there is any reason why the application cannot be allowed in this current condition, it is respectfully requested that the Examiner contact the undersigned at the number listed below to resolve any problems by Interview or Examiner's Amendment.

Respectfully submitted,

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Date: October 18, 2005

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Plank's Formula – (Quantum Mechanism)

$$E = h\upsilon = \frac{hc}{\lambda} \quad \dots\dots\dots (1)$$

E = Photon Energy
h = Plank's Constant
υ = Frequency
λ = Wavelength
c = Speed of Light

The above equation is a proven mathematical equation .

It can be represented as follows with Wavelength.

$$E[J] = \frac{h[J \cdot s]c[m/s]}{\lambda[m]} \quad \dots\dots\dots\dots\dots\dots (1)'$$

Then we convert $E[J]$ to $E[eV]$

$$E[eV] = \frac{h[J \cdot s]c[m/s]}{e[J/eV]\lambda[m]} \quad \dots\dots\dots\dots\dots\dots (2)$$

h : 6.62606876 x 10^{-34}
c : 2.99792458 x 10^{8}
e : 1.602176462 x 10^{-19}

Therefore

$$E[eV] = \frac{1240 \times 10^{-9}}{\lambda[m]} \quad \dots\dots\dots\dots\dots\dots (3)$$

$$E[eV] = \frac{1240}{\lambda[nm]} \quad \dots\dots\dots\dots\dots\dots (3)'$$

Thus

The equation (3)' is also a proven mathematical equation as well as a self-evident event.

Reference: http://www.phys.ksu.edu/gene/f_11.html

	Eg[eV]	Wavelength[nm]	E [eV]= 1240/ λ [nm]	
1	2.19	566	E [eV]= 1240/ 566[nm]	E[ev]= 2.190813
2	2.16	574	E [eV]= 1240/ 574[nm]	E[ev]= 2.160279
3	2.13	582	E [eV]= 1240/ 582[nm]	E[ev]= 2.130584
4	2.1	590	E [eV]= 1240/ 590[nm]	E[ev]= 2.101695
5	2.07	599	E [eV]= 1240/ 599[nm]	E[ev]= 2.070117
6	2.04	608	E [eV]= 1240/ 608[nm]	E[ev]= 2.039474
7	2.01	617	E [eV]= 1240/ 617[nm]	E[ev]= 2.009724
8	1.98	626	E [eV]= 1240/ 626[nm]	E[ev]= 1.980831
9	1.95	636	E [eV]= 1240/ 636[nm]	E[ev]= 1.949686
10	1.92	645	E [eV]= 1240/ 645[nm]	E[ev]= 1.922481
11	1.89	656	E [eV]= 1240/ 656[nm]	E[ev]= 1.890244

Part F: *A Closer Look at....*

EXHIBIT 2

Light and Energy

We can use two models to help us understand how light "truly" behaves: light as bullets, and light as waves.
First think of light as consisting of a bunch of bullets that turn out to have a couple of remarkable properties! In the first place, all bullets of light, from ultraviolet to infrared move through a vacuum at the same speed, "the speed of light": 186,000 miles/second = 300,000 km/sec = 1 foot/nanosecond. All types of light move at this speed through a vacuum no matter how much energy they carry. We don't live in a vacuum, of course, and in air the speeds do differ a little bit, but only a VERY little bit: red light moves about 60 km/sec slower in air than it does in a vacuum, and blue light moves about 60 km/sec slower than the red light does. This is a tiny fraction of 300,000 km/sec, so forget it for now. Light, all light, moves at 300,000 km/sec. So the energy carried by the bullets of light is not determined by their speed; energy is an independent property of the bullet. The energy per bullet distinguishes the different types of light from each other: an infrared bullet carries less energy than a red bullet, which carries less than a blue bullet, which carries less than an ultraviolet bullet, and so on.
The second remarkable property of these "bullets" is that when they pass an atom there is a small chance that they will be absorbed by the atom. They will just disappear. Really. If 100 bullets of UV A-B start at the top of the atmosphere, only 66 will make it to the ground. The other 34 will have been absorbed by the atoms in the atmosphere; they will have just vanished, leaving their energy in the atoms that have absorbed them.
A truly remarkable thing about light is that you can also think of it as a wave on water! A water wave usually moves straight ahead, just like light does. A water wave also bends a little bit when it goes around obstacles. That behavior also happens with visible light, although it is hard to see. In a dark room, aim a beam of light at the edge of a razor blade and then image the blade edge on a screen. You will see that the light has NOT gone in straight lines but has in fact gone a little bit around the blade. The image of the edge of the blade is tinged in red because the red light bends more than does the blue. Radio waves offer a more common experience of this effect. Radio reception "fades out" as you drive into a tunnel instead of suddenly cutting off, because the radio waves bend a little bit into the tunnel. So light bends around obstacles much as water waves or radio waves do. The longer the wavelength of a wave on water the more it will bend; we can color the picture of light as a wave by attributing different wavelengths to the different colors: red light bends more than blue light so it has a longer wavelength, blue light is longer than ultraviolet, and so on.
These bending effects may suggest to you that the bullet picture is wrong, that light really is more like a wave in something. But all sufficiently sensitive detectors of light find that light really does come in lumps. Light is "quantized." If light was only a wave in something then we ought to detect the energy in, say, 1/7 of the wave. But we don't. The minimum amount of energy is always just the energy in the appropriate bullet.

This fact of nature has another nice consequence; it solves a problem with the wave picture. The problem is that the picture of light as a nice continuous wave turns out to imply (using a difficult line of reasoning from 19th century physics) that any warm object, even a human body, should emit an infinite amount of energy! This energy would be in wavelengths shorter than the ultraviolet, so this impossible prediction is called the "ultraviolet catastrophe." Recognizing that light comes in lumps solves that, as Max Planck discovered in 1899.

So what is light? Is it bullets, because it comes in lumps? or waves, because it bends around stuff?

It is neither and both. Light is a unique aspect of nature. Our job is to learn how it behaves, and what it is going to do in any situation. Our current understanding of light, called the "quantum theory of light," or "quantum electrodynamics," predicts light to have just the strange sort of behaviors we have identified. Rather than delving into this complex branch of physics, we will instead describe the results of this theory and see how they help us understand how light of all types behaves and how it interacts with atoms.

Light can be thought of as a wave with a wavelength, or as a photon (the proper name of the bullet) carrying energy. The relationship between the two models of light is this:

$$\text{Energy of photon} = hc/\lambda.$$

The constants h and c are numbers that depend on the units in which we measure distance, time, and energy. "c" is the speed of light, and can be 186,000 if we're using miles and seconds, or 1 if we are using feet and nanoseconds, and so on. "h" is called "Planck's constant" and also has different values in different units. We'll pick our units below. Notice how the energy and wavelength are related. Long waves correspond to low energy photons, short waves correspond to high energy photons. Energy and wavelength are "inversely proportional."

The quantum theory also helps us understand how light is sometimes more bullet-like and sometimes more wavelike. The intuitive picture of photons as bullets that don't bend is pretty accurate for high energy photons (short wavelengths), while the wave picture is pretty accurate for long wavelengths (low energy photons). The reason is easy to see: short waves don't bend much, so the bullet picture works better, while the energy in low energy photons comes in such small lumps that it is hard to see that there are any lumps at all.

Now we can get more quantitative, decide on units to use, learn how to produce ultraviolet light, study how it interacts with matter, and so on.

Photons and Electromagnetic Waves

Photons carry energy, and waves are characterized by their wavelength. We need to settle on units for energy lumps and for wavelengths.

Wavelength is easiest: units like meters, miles, inches, and so on are familiar to you already but are too big to be convenient when talking about light and atoms. For example, a typical wavelength of visible light is 0.000000550 meters, so if we use meters we'll have to keep track of lots of zeros. We could either use exponents, and call that wavelength 550 times ten to the minus-nine meters, or invent a name for ten to the minus-nine meters, nanometers, and call that wavelength 550 nanometers. That is what we'll do. The wavelengths we discuss will be given in nanometers, abbreviated nm. Here is part of a table you have seen before: infrared light has wavelengths longer than 760 nm, visible light has wavelengths from 760 to 400 nm, and ultraviolet light has wavelengths shorter than 400 nm. A nanometer is also a good unit for talking about atoms: a hydrogen atom is about 1/10 nm across, for example, and a big atom like cesium or barium is about 6/10 nm in diameter.

Energy comes in an enormous variety of units, but all refer to the same idea. Energy in food is measured in Calories, energy from the electric company is measured in kilowatt-hours, energy from the gas company is measured in BTUs (i.e., British Thermal Units). We will find the "electron Volt," abbreviated "eV," most convenient for our purposes. When you use a 9 Volt battery in your calculator, each electron leaves the battery with 9 eV of energy and can do 9 eV of electrical work as it runs through the circuit; if your car battery is 12 Volts, then each electron can leave it with 12 eV and do 12 eV of work in the electrical wiring of your engine (actually car batteries that are rated "12 Volts" usually are not quite 12 Volts in operation; you might take up that story with a physics or electronics teacher). Most chemical reactions involve energies of a few eV per atom; the molecules in the retina of our eyes are sensitive to photons that carry between 1.6

and 3.1 eV.
Using the units we have agreed on, the quantitative connection between photons and waves looks like this:

Energy of photon (eV) = 1240 eV-nm/(wavelength in nm).

For example, orange light with a wavelength of 620 nm corresponds to photons carrying 1240 eV-nm/620 nm = 2 eV.

Practice with the numbers.
1. What is the wavelength of a 4 eV UV-B photon?
2. What is the wavelength of a 13.6 eV photon? (This is the energy needed to ionize a hydrogen atom.)
3. What is the wavelength of a 4000 eV photon? (This is an X-ray.)
4. What is the energy of a 500 nm photon? (More photons leave the sun with approximately this wavelength than any other.)
5. What is the energy of a photon whose wavelength is the size of a hydrogen atom?
6. What is the energy of a photon whose wavelength is the size of a barium atom?

How Photons Are Produced and Absorbed

Photons are produced or absorbed when electrons "change their state," that is, whenever electrons change their speed, or their direction, or their arrangement in an atom or molecule. For example, at a radio station the electrons are pushed up and down the antenna. They are constantly changing their speed and their direction of motion. As the electrons move they emit photons, radiating electromagnetic waves. The frequency assigned to a radio station indicates the number of times the electrons rush back and forth each second: 101.5 kiloHertz means that the electrons are shoved back and forth 101,500 times each second, emitting 101,500 waves each second, too. Whenever electrons rearrange themselves in an atom or molecule they emit or absorb photons. The phrase "arrangement of electrons" is a bit misleading, somewhat like calling photons "bullets" is misleading. What the phrase means is roughly "where the electron is likely to be found," or "pattern of probability." Only certain of these arrangements or patterns are allowed; they are "quantized." This is best explained by some examples.
An atom of hydrogen allows a whole series of arrangements or patterns. In chemistry and physics books these patterns are called "orbitals" and given names: 1s, 2s, 2p and so on. If the electron is in the 1s orbital, the probability is about 1/3 that the electron will be found within a small sphere just 0.1 nm across; in the 2s the electron is most likely to be about 0.2 nm from the center of the atom; in the 2p the electron is most likely to be found somewhere in a dumbbell-shaped region about 0.2 nm in radius; and so on. Only certain patterns of probability are allowed. No pattern on the list allows the electron to be found in a little sphere 0.3 nm across.

Electrons in molecules also have allowed patterns, places they are likely to be. For example, in a molecule of hydrogen, H2, there are two electrons, both likely to be found between the two nuclei. Their negative charge holds the two positive nuclei together to make the molecule. In H2O some electrons are concentrated close to the oxygen nucleus, but one pair is most likely to be found in between the O and one of the H nuclei, and another pair between the O and the other H. Again the negative charge of these localized electrons holds the molecule together. These electron arrangements are the source of the "chemical bond." The atoms in DNA are also held in place by chemical bonds; electrons stay between the atoms and hold them together because the electrons

don't have an allowed pattern that permits them to slip away.

Each arrangement has a certain energy. Low energy arrangements have electrons that are likely to be close to the nuclei, while those in which the electrons are likely to be farther away have higher energy. If an electron changes from a high energy arrangement to a lower energy arrangement it emits a photon that carries away the lost energy. For example, when a hydrogen atom changes from the 2p to the 1s arrangement, the atom loses 10.2 eV and thus emits a photon carrying 10.2 eV. 3s,3p 1.5 eV 2s,2p 3.4 eV 10.2 eV photon emitted 1s 13.6 eV Now we can describe how ultraviolet light is produced in a fluorescent light bulb, and also why, in spite of this, they are safe ordinary lighting.
When a fluorescent light is turned on, electrons start racing from one end of the tube to the other. On the way they bump into atoms, often of mercury, giving them enough energy to allow the electrons to rearrange into a higher energy pattern. Electrons in the high energy pattern in mercury quickly rearrange into the lower energy pattern, losing about 4.9 eV and so emitting a 4.9 eV photon. Now our eyes are not sensitive to photons with this much energy and they are UV-C photons besides, which are very bad for our health! But the high energy photons are absorbed by atoms in the white powder that coats the inside of the bulb. This powder is a mixture of various materials that "fluoresce," that absorb the high energy photons and emit lower energy photons. Here is how they manage this trick. Between the lowest energy arrangement of electrons in the fluorescent material and the 4.9 eV pattern there are several other energy levels. When the atom rearranges it usually does not emit a 4.9 eV photon but jumps to one of the intermediate levels, losing less energy than 4.9 eV. In fact, a whole range of energies is emitted that gives the impression of "white light," although the actual distribution of colors is not the same as sunlight. Remember that "photon energy" corresponds to "color."

Click here for figure 1

Click here for figure 2

4.9 eV state visible energy photon other levels ground state More About Energy Energy is what it takes to do work, and our idea of work grows out of everyday experiences: you do more work when you pull a heavy load than a light load; you do more work if you pull it a long distance than a short distance. The more work you do, the more energy you need.
Molecules do work, too. It takes energy to pull atoms from one place in a molecule into a new position, or to pull an atom out of one molecule and into another. The energy first goes into rearranging the electrons into a new pattern. In the new pattern the electrical forces are a bit different so the atoms start shifting around in response to these changed forces. (By the way, the electrical forces between atoms that result from the quantized patterns of the electrons are sometimes called "chemical forces.") Since the first step in molecular energy transactions is electron rearrangement, let's look at the amount of energy this involves in a number of typical situations.

Rearranging the electrons that form a chemical bond typically involves a few tenths of an eV up to a few eV. For example, to rearrange the electrons in the H2 molecule into two H atoms requires about 4.5 eV; to form the bonds in a molecule of ATP requires about 0.3 eV. These are typical energy requirements for many chemical reactions, all of which involve the rearranging of some outer electrons of atoms. If the reaction goes the other way, of course the energy is released. For example, if 2H turns into H2 then 4.5 eV are released when each molecule is formed. In our bodies, and in yeast cells, the bonds in a molecule of ATP are broken to release 0.3 eV of energy the cell needs to do its work.
Here are some examples involving different amounts of energy.

a) Hammering a nail, one blow: This involves lots of eV, around 10^19, but this energy is spread out through LOTS of atoms, maybe 1023 or so. Thus the energy per atom is very small, about 10-4 eV, too small to cause any chemical reaction. Instead, the atoms just shift position a little bit, their electrons remaining as they were.
b) Fire: This involves lots of different chemical reactions, but a very important one is oxygen and carbon combining to form carbon dioxide. Each time this happens 4.1 eV of energy are released. Much of this energy goes to make the molecules move faster, that is, to make the gas hot.
c) Batteries: When an atom gains or loses some of its outer electrons, its energy changes. This is called its "electromotive force" in chemistry and physics books, and is typically between 0.1 and 3 eV. For example, at the positive electrode of a lead storage battery, the lead gains two electrons and at the negative electrode it loses two electrons; the difference in energy of the two processes is just over 2 eV. (This depends a bit on the concentration of the lead ions, but that is a more advanced topic. We want to leave a little for your chemistry and physics courses to explain!)

Relationships Between Various Units of Energy

We've been using electron Volts (eV) as the unit of energy because it takes around an eV to significantly affect most molecules, and we're interested in individual molecules. But when the question involves lots of molecules, eV is too small a unit because no one wants to always deal with huge powers of ten, or, worse yet, long strings of zeros in numbers. For example, the energy in a gallon of gasoline is about 1 1027 eV = 1,000,000,000,000,000,000,000,000,000 eV. When we talk about the energy content of ordinary amounts of matter we need more reasonable units than eV: Joule, kWhr, Cal, BTU, and others. Although we will not usually need these other units in discussing the experiments on ultraviolet light, we've included some examples of these other units below, followed by a table showing the connections between them.
We are accustomed to electrical energy. If we use a 100 Watt light bulb for 10 hours then we have used 1000 Watt-hours = 1 kW-hr of electrical energy. A typical household uses around 5000 kWhr each month. The energy consumed by the power plant to provide this to us is about three times this much because generating and distributing electricity is not very efficient.
The food we eat and oxygen we breathe supplies us with energy we need to carry out the business of life. Oxygen combines with the molecules from food to release the energy that our body uses in a process called "cellular respiration." The typical unit of energy in food is the Calorie, sometimes called a "kilocalorie" or "large calorie." People need a few thousand of these Calories each day. The Calories in food really are the same sort of energy as the energy in gasoline or coal. We could burn potatoes to run power plants, and getting 2000 Calories from them would produce the same electrical energy that getting 2000 Calories from burning coal yields. The amazing feat of our metabolism is that it can extract the Calories from potatoes by oxidation at body temperature instead of at 700 degrees as done in the power plant!
The energy in coal, oil, and natural gas is often expressed in BTUs, an acronym for British Thermal Unit, the unit in which national energy debates are often conducted. Ten gallons of gasoline contain roughly a million BTUs. The United States uses around 80 quadrillion BTUs per year. That is 80 1015 BTUs, or "80 quads." The energy content of one barrel of oil is about 6 million BTU, and we import around 15 quads of oil. You can figure out how many billion (yes, billion) barrels of oil we import each year.
Here are the approximate conversions between the units we've illustrated above (and some we haven't):

Energy Conversion Table

Table 1: Energy Conversion Table

	eV	Cal	kW-hr	BTU	J	ft-lb
1 eV =	1	1.4 10-23	1.7 10-26	5.9 10-23	6.2 10-20	4.6 10-^20
1 Cal =	7.0 1022	1	1.2 10-3	4	4200	3100
1 kW-hr =	5.8 1025	860	1	3400	3.6 106	2.7 106
1 BTU =	1.7 1022	0.25	2.9 10-4	1	1100	780
1 J =	1.6 1019	2.4 10-4	2.8 10-7	9.5 10-4	1	0.74
1 ft-lb =	2.2 1019	3.4 10-4	3.9 10-7	1.3 10-3	1.4	1

Ordinary amounts of matter contain many eVs of energy because ordinary amounts of matter contain so many atoms. The basic unit for number of atoms is the "mole," or "Avogadro's number." If you have 6 1023 atoms of a substance then you have one mole of the stuff. Since different atoms have different weights you can measure out a mole by weighing the stuff. One gram of hydrogen contains 6 1023 atoms of hydrogen; it is one mole of hydrogen. Four grams of helium make up one mole of helium and so it contains 6 1023 atoms of helium; 12 grams of carbon equal one mole of carbon; and so on. If you check the atomic weights (which are on most periodic tables) of hydrogen, helium and carbon you will quickly get the point: whatever the atomic weight of an element is, that many grams of the element contains one mole of atoms.
Since a chemical reaction involves around one eV per atom we should expect one mole of atoms to produce around 6 1023 eV. From the table above we see that this is about 23 Calories. Sure enough, the typical "heats of formation" studied in chemistry are in this range for simple molecules.

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